Answers to Practice Test Questions 5
Electron Configurations and Periodic Trends

1.

Symbol	Electronic Configuration	Number of Valence Electrons
Se	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^4$	6
Fe	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$	8
S^{2-}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	8

2.

Electron Configuration	Could describe a ground state atom? Circle yes or no.	If no, why not? If yes, name the element.
$1s^2\ 2s^2\ 2p^2$	YES NO	Carbon (C)
$1s^1\ 2s^2$	YES NO	This atom is an excited state. Ground state for three electrons would be $1s^2\ 2s^1$.
$1s^2\ 2s^2\ 2p^5$	YES NO	Fluorine (F)
$1s^2\ 2s^2\ 3s^2$	YES NO	This atom is an excited state. Ground state for six electrons would be $1s^2\ 2s^2\ 2p^2$.

3.

(a) $1s^2\ 2s^2\ 2p^5$

(b) any electron configuration with 9 electrons that is not $1s^2\ 2s^2\ 2p^5$ and does not violate the Pauli exclusion principle

e.g. $1s^1\ 2s^2\ 2p^6$

4.

(a) [Ar] $4s^1\ 3d^{10}$

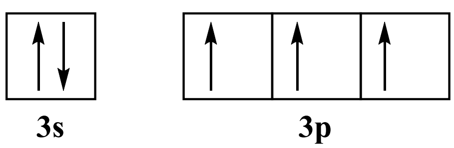
(b) [Ar] $3d^{10}$

(c) Cu^+ is diamagnetic. All its electrons are spin-paired.

5.

(a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

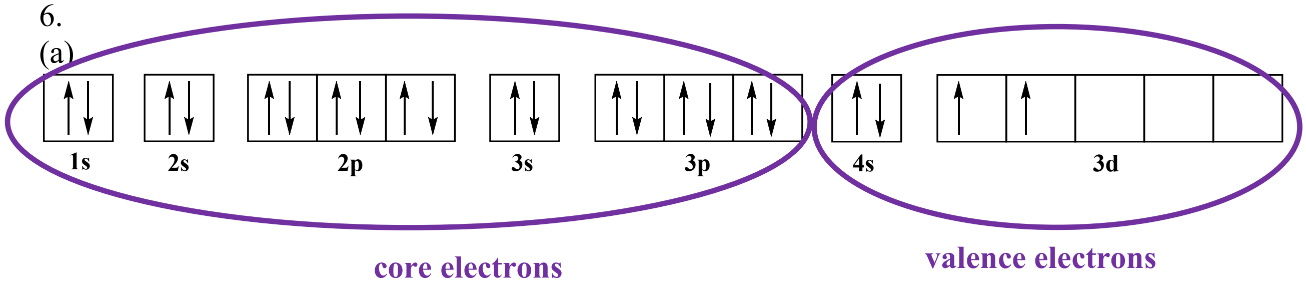
(b)



(c) paramagnetic

6.

(a)



(b) *see above*

(c) paramagnetic

(d) Ti^{+4}

Removing four electrons from Ti gives Ti^{4+} with a noble gas electron configuration.

Note that the 4s electrons are removed before the 3d electrons, so the electron configuration of Ti^{2+} would be [Ar]$3d^2$ <u>not</u> [Ar]$4s^2$.

7.

(a) [Kr] $5s^2\ 4d^{10}\ 5p^2$

(b) *Any <u>one</u> of the following six answers is acceptable:*

$n = 5 \quad l = 1 \quad m_l = +1 \quad m_s = +½$

$n = 5 \quad l = 1 \quad m_l = +1 \quad m_s = -½$

$n = 5 \quad l = 1 \quad m_l = 0 \quad m_s = +½$

$n = 5 \quad l = 1 \quad m_l = 0 \quad m_s = -½$

$n = 5 \quad l = 1 \quad m_l = -1 \quad m_s = +½$

$n = 5 \quad l = 1 \quad m_l = -1 \quad m_s = -½$

8.

(a) If bromine forms an ion, its charge is _-1_.

(b) If aluminium forms an ion, its charge is _+3_.

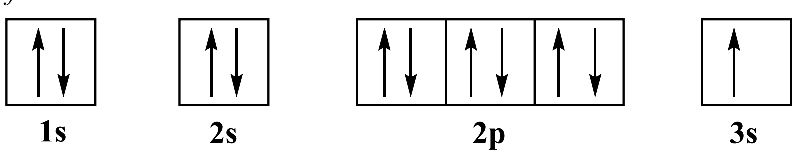
(c) If thallium (Tl) forms an ion, its charge is either _+1_ or _+3_.

9.

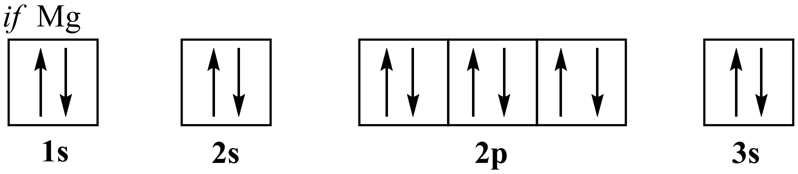
(a) *any two of:* $\mathbf{Na^+}$ $\mathbf{Mg^{2+}}$ $\mathbf{Al^{3+}}$

(b)

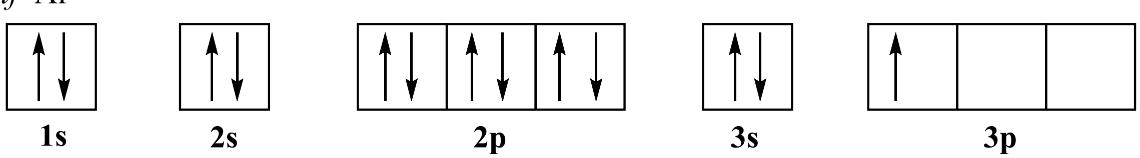
if Na



if Mg



if Al



10.

(a) [Kr] $5s^2$ $4d^2$

(b) [Ar] $4s^2$ $3d^{10}$ $4p^6$

(c) Zr

Zr has electrons in the $n = 5$ shell while Zr^{4+} does not. These electrons are more shielded than the valence electrons in Zr^{4+}, so they are less strongly attracted to the nucleus. Thus, Zr has a larger radius than Zr^{4+}.

11.

(a) F $1s^2$ $2s^2$ $2p^5$ F^- $1s^2$ $2s^2$ $2p^6$

(b) F

F and F^- have the same number of protons, but F^- has one more electron. As such, the valence electrons are more shielded from the nuclear charge in F^- than they are in F. Therefore, the valence electrons in F feel a higher effective nuclear charge.

12. Fe^{2+} and Fe^{3+} both have 26 protons in their nuclei; however, Fe^{2+} has 24 electrons ([Ar] $3d^6$) while Fe^{3+} has 23 electrons ([Ar] $3d^5$).

Because Fe^{2+} has more electrons, its valence electrons are more shielded from the positive charge of the nucleus and therefore feel a weaker effective nuclear charge. Thus, they are attracted less strongly toward the nucleus, giving Fe^{2+} a larger ionic radius than Fe^{3+}.

13. Electron affinity is defined as the energy released when an electron is added to a neutral atom in the gas phase.

F has a larger electron affinity than Ne.

It is much more favourable to add an electron to F (giving F^-) than it is to add an electron to Ne (giving Ne^-). To add an extra electron to Ne, it would be necessary to begin a new shell that would be strongly shielded from the nuclear charge so the new electron would not be strongly attracted to the nucleus. On the other hand, the extra electron added to F is added to the valence shell, giving a noble gas electron configuration of [He] $2s^2$ $2p^6$.

14. Ca

Ca and Ca^{2+} have the same number of protons; however, Ca^{2+} has two fewer electrons:

Ca [Ar] $4s^2$ Ca^{2+} [Ne] $3s^2$ $3p^6$

The valence electrons in Ca are in the $n = 4$ shell whereas the valence electrons in Ca^{2+} are in the $n = 3$ shell (which is closer to the nucleus).

15.

Description	Element
An element with a larger first ionization energy than chlorine.	helium(He), fluorine (F), neon (Ne), *or* argon (Ar)*
The Group 13 element with the smallest atomic radius.	boron (B)
The element which has the following electron configuration: [Ar] $4s^2\ 3d^8$	nickel (Ni)
An element whose trication (cation with +3 charge) has the same number of electrons as vanadium (V).	iron (Fe)
An element whose dianion (anion with -2 charge) has the same electron configuration as krypton (Kr).	selenium (Se)
The first element to have an electron in a *d* orbital when in the ground state. *("first" = "with smallest atomic number")*	scandium (Sc)

* *Nitrogen (N), oxygen (O) and krypton (Kr) also meet this description; however, that is not obvious from the periodic trends and you are* ***not*** *expected to have memorized any of the elements' ionization energies.*

16.

(a) Se [Ar] $4s^2\ 3d^{10}\ 4p^4$ Kr [Ar] $4s^2\ 3d^{10}\ 4p^6$

electron configurations must show the valence electrons, so Kr shouldn't just be [Kr]

(b) Kr

It is more difficult to remove a valence electron from Kr than it is from Se. This is because the valence electrons in Kr feel a greater effective nuclear charge (since it has two more protons in its nucleus and its extra two electrons do not fully shield their charge).

Your answer must clearly show that you understand the meaning of the term "first ionization energy".

(c) Se

It is easier to add an electron to Se than it is to Kr. The electron added to Se goes in the 4p subshell while an electron added to Kr would go in the 5s orbital, beginning a new electron shell. Thus, the electron added to Kr would feel very little attraction to the nucleus because it would be shielded by all of the electrons in the first four shells. In fact, it actually requires energy to add an electron to Kr whereas energy is released when an electron is added to Se.

Your answer must clearly show that you understand the meaning of the term "electron affinity".

17. Mg^{2+}

All three ions have 10 electrons, however, Mg^{2+} has 12 protons, Na^{+} has 11 protons and F^{-} has 9 protons. Because Mg^{2+} has the most protons, its electrons experience a higher effective nuclear charge and are more strongly attracted to the nucleus, reducing the radius of the ion.

18. If an atom has electrons in both the 2s and 2p orbitals, they will shield each other from the nucleus; however, the 2s electrons will shield the 2p electrons more than the 2p electrons will shield the 2s electrons.

This is due to the fact that the 2s electrons spend a small amount of time very close to the nucleus (the 2p electrons spend much less time this close). This phenomenon is called penetration.

Recall when radial nodes were shown in class. The 2s orbital has one radial node while the 2p orbital has no radial nodes. The phenomenon described above is a consequence of this difference.

19. The electron configuration for boron is $1s^2\ 2s^2\ 2p^1$

The first electron ionized from boron comes from a 2p orbital. The second electron comes from the 2s orbital. Since the 2s orbital is lower in energy than the 2p orbitals, the second ionization energy is significantly larger.

This is in addition to the usual increase from first ionization energy to second ionization energy caused by reduced electron shielding when the first electron is removed.

20.

(a) $Be \rightarrow Be^{+} + e^{-}$

(b) $Be^{3+} \rightarrow Be^{4+} + e^{-}$

(c) The fourth ionization energy is larger.

Removing an electron from Be^{3+} requires more energy than removing an electron from Be because the electron in Be^{3+} feels the whole +4 charge from the nucleus whereas the valence electrons in Be are shielded by each other as well as the two core electrons.

(d) The fourth ionization energy of beryllium is easier to calculate because Be^{3+} only has one electron so there is only one force of attraction to calculate since we don't have to factor in electron-electron repulsion (or shielding).

(e) The ionization energy for ground state Be^{3+} (i.e. the fourth ionization energy of Be) is equal to the energy difference between an electron with n = 1 and an electron with $n = \infty$ (the ionization limit). The energy for an electron with $n = \infty$ is 0 J.

Be^{3+} has four protons so $Z = 4$.

Step 1: Calculate the energy for Be^{3+} with n = 1

$$E_{n=1} = -R_H \frac{Z^2}{n^2} = -R_H \frac{(4)^2}{(1)^2} = -16R_H = -3.487795 \times 10^{-17} J$$

Step 2: Calculate the ionization energy for ground state Be^{3+}

$$E_i = E_{n=\infty} - E_{n=1} = (0J) - (-3.487795 \times 10^{-17} J) = 3.487795 \times 10^{-17} J$$

Step 3: Convert the ionization energy from J/atom to kJ/mol

$$E_i = 3.487795 \times 10^{-17} \frac{J}{atom} \times \frac{1kJ}{1000J} \times \frac{6.022141 \times 10^{23} atoms}{1mol}$$

$$E_i = 21003.99 \frac{kJ}{mol} = 2.100399 \times 10^4 \frac{kJ}{mol}$$

Step 4: Check your work

Does your answer seem reasonable? Are sig. fig. correct?

The ionization energy (per atom) has a similar order of magnitude to R_H.

The ionization energy (per mole) is much larger.

21.

(a) oxygen (O)

6 valence electrons = group 6 or 16

The only element in either group likely to have a stable isotope with only 8 neutrons (i.e. atomic mass of approximately 2 × 8 u = 16 u) is oxygen. Oxygen has 8 protons and an average atomic mass of 15.9994 u.

(b) beryllium (Be)

If half the electrons in an atom are valence electrons then half of them are core electrons.

There are only a few possible values for the number of core electrons. Any element in period 1 has 0 core electrons; any element in period 2 has 2 core electrons ($1s^2$); any element in period 3 has 10 core electrons ($1s^2 2s^2 2p^6$); any element in period 4 has 18 core electrons ($1s^2 2s^2 2p^6 3s^2 3p^6$); etc.

Beryllium has 2 core electrons plus 2 valence electrons. Note that there is no other element that meets these criteria (calcium has 20 electrons, but 18 are core – not 10).

(c) carbon (C)

2 core electrons and 4 valence electrons

see logic for part (b); for any other valid number of core electrons in a ground state atom, it is not possible to have twice as many valence electrons.

(d)

lithium (Li)	2 core electrons + 1 valence electron
phosphorus (P)	10 core electrons + 5 valence electrons
cobalt (Co)	18 core electrons + 9 valence electrons

see logic for part (b)

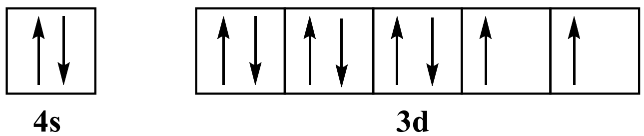
note that xenon (Xe) has 8 valence electrons – not 18!

22.

(a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^8$

(b) $[Ar]\ 4s^2\ 3d^8$

(c)



(d)

Electron	***n***	***l***	m_ℓ	m_s
1 (4s)	4	0	0	+ ½
2 (4s)	4	0	0	– ½
3 (3d)	3	2	+2	+ ½
4 (3d)	3	2	+1	+ ½
5 (3d)	3	2	0	+ ½
6 (3d)	3	2	-1	+ ½
7 (3d)	3	2	-2	+ ½
8 (3d)	3	2	+2	– ½
9 (3d)	3	2	+1	– ½
10 (3d)	3	2	0	– ½
11				
12				

Each row corresponds to one electron. For clarity, bold lines have been used to separate the rows into three sets – the two 4s electrons, the five 3d electrons with the same spin as each other (shown with arrows pointing up), and the other three 3d electrons (shown with arrows pointing down). Since there were only ten valence electrons, the bottom two rows of the table were not needed.